APPENDIX

From Chambers Science and Technology Dictionary

metallic bond *(Chem.)*. In metals, the valence electrons are not even approximately localized in discrete covalent bonds, but are delocalized and interact with an indefinite number of atomic nuclei. This gives rise both to the opacity and lustre of metals and to their electrical conductivity.

From The Columbia Encyclopedia, Sixth Edition, 2001

Metallic and Hydrogen Bonds

Unlike the ionic and covalent bonds, which are found in a great variety of molecules, the metallic and hydrogen bonds are highly specialized. The metallic bond is responsible for the crystalline structure of pure metals. This bond cannot be ionic because all the atoms are identical, nor can it be covalent, in the ordinary sense, because there are too few valence electrons to be shared in pairs among neighboring atoms. Instead, the valence electrons are shared collectively by all the atoms in the crystal. The electrons behave like a free gas moving within the lattice of fixed, positive ionic cores. The extreme mobility of the electrons in a metal explains its high thermal and electrical conductivity.

Hydrogen bonding is a strong electrostatic attraction between two independent polar molecules, i.e., molecules in which the charges are unevenly distributed, usually containing nitrogen, oxygen, or fluorine. These elements have strong electron-attracting power, and the hydrogen atom serves as a bridge between them. The hydrogen bond, which plays an important role in molecular biology, is much weaker than the ionic or covalent bonds. It is responsible for the structure of ice.